Chemistry B11
Chapter 4
Chemical bonds

<u>Octet rule</u>: when undergoing chemical reaction, atoms of group 1A-7A elements tend to gain, lose, or share sufficient electrons to achieve an electron configuration having eight valance electrons.

Note: for the main-group elements (1A-7A), the first shell can have a maximum 2 electrons (for hydrogen and helium) and other shells 8 electrons.

Metal (usually): loses one, two or three electrons and in losing electrons, the atom becomes a positively charged ion called a **<u>Cation</u>** (Na^+, Ca^{2+}, Mg^{2+}).

Nonmetal (usually): gains one, two or three electrons and in gaining electrons, the atom becomes a negatively charged ion called an **<u>Anion</u>** (Cl^-, O^{2-}, S^{2-}).

Note: the octet rule is not perfect for two reasons:

1. Ions of period 1 and 2 elements with charges greater than +2 are unstable (B^{3+}, C^{4+} and C^{4-} don't exist because they are unstable). It is far too large a charge for an ion of these period elements.
2. The octet rule does not apply to the transition and the inner transition elements (groups 1B-7B).

<u>Naming Monatomic cations</u>: name of metal + "ion"

Na^+ Sodium ion Ba^{2+} Barium ion Al^{3+} Aluminium ion

If elements have more than one type of cation (most transition and inner transition elements), we show the charge with the Roman numeral immediately following the name of the metal (for **<u>Systematic name or IUPAC</u>** (*International Union of Pure and Applied Chemistry*)). We can also use the suffix "-ous" to show the smaller charge and the suffix "-ic" to show the larger charge (for **<u>Common name</u>**).

Ion	Systematic name	Common name
Cu^+	Copper(I) ion	Cuprous ion
Cu^{2+}	Copper(II) ion	Cupric ion
Fe^{2+}	Iron(II) ion	Ferrous ion
Fe^{3+}	Iron(III) ion	Ferric ion
Hg^+	Mercury(I) ion	Mercurous ion
Hg^{2+}	Mercury(II) ion	Mercuric ion
Sn^{2+}	Tin(II) ion	Stannous ion
Sn^{4+}	Tin(IV) ion	Stannic ion
Cr^{2+}	Chromium(II)	Chromous ion
Cr^{3+}	Chromium(III)	Chromic ion
Co^{2+}	Cobalt(II)	Cobaltous ion
Co^{3+}	Cobalt(III)	Cobaltic ion
Pb^{2+}	Lead(II)	Plumbous ion
Pb^{4+}	Lead(IV)	Plumbic ion

Naming Monatomic anions: we add "-ide" to the step part of the name.

Anion	Stem name	Anion name
F^-	fluor	Fluoride
Cl^-	chlor	Chloride
Br^-	brom	Bromide
I^-	iod	Iodide
S^{2-}	sulf	Sulfide
O^{2-}	ox	Oxide
N^{3-}	nitr	Nitride
P^{3-}	phosph	Phosphide

Naming Polyatomic ions: we use the prefixes "di-", "tri-" and so forth to show the presence of more than one hydrogen (the prefix "bi" is used to show the presence of one hydrogen).

Polyatomic ion	Name	Polyatomic ion	Name
NH_4^+	Ammonium	SO_3^{2-}	Sulfite
OH^-	Hydroxide	SO_4^{2-}	Sulfate
NO_2^-	Nitrite	HSO_3^-	Hydrogen sulfite (bisulfite)
NO_3^-	Nitrate	HSO_4^-	Hydrogen Sulfate (bisifate)
CH_3COO^-	Acetate	PO_3^{3-}	Phosphite
CN^-	Cyanide	PO_4^{3-}	Phosphate
MnO_4^-	Permanganate	HPO_4^{2-}	Hydrogen phosphate
CrO_4^{2-}	Chromate	$H_2PO_4^-$	Dihydrogen phosphate
$Cr_2O_7^{2-}$	Dichromate	ClO^-	Hypochlorite
CO_3^{2-}	Carbonate	ClO_2^-	Chlorite
HCO_3^-	Hydrogen carbonate (bicarbonate)	ClO_3^-	Chlorate
		ClO_4^-	Perchlorate

Ionic bonds: ionic bonds usually form between a metal and a nonmetal. In ionic bonding, electrons are completely transferred from one atom to another. In the process of either losing or gaining negatively charged electrons, the reacting atoms form ions. The oppositely charged ions are attracted to each other by electrostatic forces, which are the basis of the ionic bond.

$$Na\ (1s^2\ 2s^2\ 2p^6\ 3s^1) + Cl\ (1s^2\ 2s^2\ 2p^6\ 3s^1\ 3p^5) \rightarrow Na^+\ (1s^2\ 2s^2\ 2p^6) + Cl^-\ (1s^2\ 2s^2\ 2p^6\ 3s^1\ 3p^6)$$

Note: there are enormous differences between the chemical and physical properties of an atom and those of its ion(s). For example sodium is a soft metal and it reacts violently with water. Chlorine is a gas and it is very unstable and reactive. Both sodium and chlorine are poisonous. However, NaCl (common table salt made up of Na^+ and Cl^-) is quite stable and unreactive.

Note: the charge of a monatomic ion can be predicted from the periodic table:

Group	# Valence e^-	# Gained/Lost	Charge on Ion
1A	1	1	+1
2A	2	2	+2
3A	3	3	+3
5A	5	3	-3
6A	6	2	-2
7A	7	1	-1

Note: matters are electrically neutral (uncharged). The total number of positive charges must equal the total number of negative charges. The subscripts in the formulas for ionic compounds represent the ratio of the ions.

$Na^+ \quad Cl^- \rightarrow NaCl \qquad Ca^{2+} \quad Cl^- \rightarrow CaCl_2 \qquad Al^{3+} \quad S^{2-} \rightarrow Al_2S_3$

$Ba^{2+} \quad O^{2-} \rightarrow Ba_2O_2$ we must reduce to lowest terms: BaO

<u>Naming binary ionic compounds</u>: Name of cation (metal) + name of anion

Note: We generally ignore subscripts in naming binary ionic compounds.

Note: many transition metals form more than one positive ion. We use Roman numerals in the name to show their charges.

$NaCl$ Sodium chloride $\quad$ CaO Calcium oxide $\quad$ $AlCl_3$ Aluminium chloride

CuO Copper(II) oxide (cupric oxide) $\quad$ $FeCl_2$ Iron(II) chloride (Ferrous chloride)

$MgCO_3$ Magnesium carbonate $\quad$ $NaOH$ Sodium hydroxide

<u>Covalent bonds</u>: covalent bonds usually form between two nonmetals or a metalloid and a nonmetal. In covalent bonds, the atoms share one or more pairs of electrons (by using their valence electrons) between each other to obtain a filled valence shell.

H-H $\quad$ C-H $\quad$ O-H

<u>Electronegativity</u>: electronegativity is measure of an atom's attraction for the electrons it shares in a chemical bond with another atom. The electronegativity shows us how tightly an atom holds the electrons that it shares with another atom.

Note: electronegativity generally increases from left to right across a row of the Periodic Table.

Note: electronegativity generally increases from bottom to top within a column of the Periodic Table.

Note: when the electronegativity increases, the ionisation energy increases too.

We classify covalent bonds into two categories:

1. Nonpolar covalent bonds: bonding with an equal sharing of electrons.

H-H   Cl-Cl   N-N

2. Polar covalent bonds: bonding with an unequal sharing of electrons. The number of shared electrons depends on the number of electrons needed to complete the octet.

O-H   O-F   N-H

The less electronegative atom has a lesser fraction of the shared electrons and obtains a partial positive charge (δ+). The more electronegative atom gains a greater fraction of the shared electrons and obtains a partial negative charge (δ-). This separation of charge produces a **dipole** (two poles). We show a bond dipole by an arrow, with the head of the arrow near the negative end of the dipole and a cross on the tail of the arrow.

$$\overset{\delta+}{\mathbf{H}} - \overset{\delta-}{\mathbf{Cl}}$$

$$\longmapsto$$

Note: we use the following table to find the type of a chemical bond:

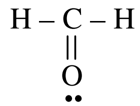
Electronegativity difference between bonded atoms	Type of bond
less than 0.5	nonpolar covalent
0.5 to 1.9	polar covalent
greater than 1.9	Ionic

Examples:

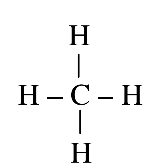
H-Cl	3.0 - 2.1 = 0.9	Polar covalent bond
C-H	2.5 - 2.1 = 0.4	Nonpolar covalent bond
Zn-O	3.5 - 1.6 = 1.9	Ionic bond

Covalent compounds: the most common method to show the covalent compound is Lewis Structure. First, we should determine the number of valence electrons in the molecules. Then, we connect the atoms by single bonds. Finally, we arrange the remaining electrons so that each atom has a complete outer shell.

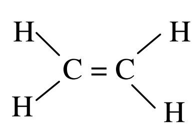
CH_2O



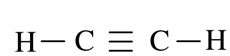
CH_4



C_2H_4



C_2H_2



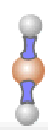
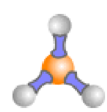
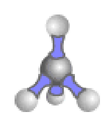
<u>Naming binary covalent compounds</u>: name the less electronegative element (the first element in the formula) + name the more electronegative element (the second element in the formula) + adding "-ide" to the stem part of the name. We use the prefixes mono-, di-, tri-, tetra-, penta-, hexa-, hepta-, octa-, nona-, deca- to show the number of atoms of each element.

Note: the prefix "mono-" is omitted for the first atom named and it is rarely used with the second atom.

Note: we drop "a" when following a vowel.

SO_3	Sulfur trioxide	PCl_5	Phosphorus pentachloride
N_2O_4	Dinitrogen tetroxide	OF_2	Oxygen difluoride

<u>Bond angle (VSEPR model)</u>: the angle between two atoms bonded to a central atom. According to VSEPR model, the valance electrons of an atom may be involved in the formation of single, double, or triple bonds, or they may be unshared. Each combination creates a negatively charged region of electron density around a nucleus. Because like charges repel each other, these regions of electrons go as far away as possible from each other. Depending on the number of these regions of electrons around a central atom, we can have three different possibilities:

Angles between *bonding* pairs	180°	120°	109.5°
Name of shape	Linear	Trigonal Planar	Tetrahedral
			

Note: the presence of the unshared pairs of electron can change these angles. Because, the unshared pairs of electrons repel adjacent electron pairs more strongly than bonding pairs repel one another.

<u>Polarity</u>: a molecule will be polar if it has two properties at the same time:

1. It has polar bonds.
2. Its centers of partial positive charge and partial negative charge lie at different places within the molecule.

Note: in order to find the polarity of a molecule, we should:

1. Draw the Lewis Structure.
2. Determine the bond angle and the shape of the molecule.
3. Predict the electronegativity of which atom is higher than others.
4. Indicate partial positive charge and partial negative charge for each atom and draw the dipole signs.

5. If the centers of partial positive charge and partial negative charge lie at the different places within the molecule, molecule would be polar. Otherwise, molecule is nonpolar.